

Acid Base

1



$0.14 \text{ M HOCl} \times 40.0 \text{ mL} \times \frac{1 \text{ mol NaOH}}{1 \text{ mol HOCl}} \times \frac{1 \text{ L}}{0.56 \text{ mol NaOH}} = \boxed{10.0 \text{ mL NaOH}}$



$K_a = \frac{[\text{H}_3\text{O}^+][\text{OCl}^-]}{[\text{HOCl}]} = \frac{x \cdot x}{0.14 \text{ M}} = 3.2 \times 10^{-6}$

c

I	0.14 M	0	0
C	-x	+x	+x
E	0.14 M - x	x	x

Assume $0.14 \text{ M} \gg x$
 Then $0.14 \text{ M} - x \approx 0.14 \text{ M}$
 $[H^+] = x = 6.7 \times 10^{-4} \text{ M}$
 $5\% > 6.7 \times 10^{-4} \text{ ok}$

$\text{pH} = -\log(6.7 \times 10^{-4} \text{ M}) = \boxed{3.2}$

d $\% \text{ ionization} = \frac{6.7 \times 10^{-4} \text{ M}}{0.14 \text{ M}} \times 100 = 0.48\%$

e $3.8 \text{ g NaOCl} \times \frac{1 \text{ mol NaOCl}}{74.5 \text{ g NaOCl}} = 0.051 \text{ mol NaOCl}$
 $\left(\frac{100 \text{ mL} \times 1 \text{ L}}{1000 \text{ mL}} \right) = 0.51 \text{ M NaOCl}$

$\text{HOCl} \rightleftharpoons \text{H}^+ + \text{OCl}^-$

I	0.14 M	0	0.51 M
C	-x	+x	+x
E	0.14 M - x	x	0.51 M + x

Assume $0.14 \text{ M} \gg x$
 $0.14 \text{ M} - x \approx 0.14 \text{ M}$

$K_a = \frac{[H^+][\text{OCl}^-]}{[\text{HOCl}]} = 3.2 \times 10^{-6}$
 $[H^+] = 8.8 \times 10^{-7} \text{ M}$

$\% = \frac{8.8 \times 10^{-7} \text{ M}}{0.14 \text{ M}} \times 100 = \boxed{6.3 \times 10^{-4} \%}$



$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{3.2 \times 10^{-6}} = \boxed{3.1 \times 10^{-9}}$ (P)

From a

I	5.6 mmol	5.6 mmol	0	0
C	-5.6	-5.6	+x	+x
E	0	0	x	x

Assume $0.11 \text{ M} \gg x$
 $\left(\frac{5.6 \text{ mmol}}{40 \text{ mL} + 60 \text{ mL}} \right) = \boxed{0.11 \text{ M OCl}^-}$ (g)

$K_b = \frac{x^2}{(0.11 \text{ M})} = 3.1 \times 10^{-9}$

$[\text{OH}^-] = x = 1.9 \times 10^{-5}$

$\text{pH} = 14 - (-\log 1.9 \times 10^{-5}) = \boxed{9.27}$ (h)

i $0.56M NaOH \times 5.0 mL = 2.8 mmol NaOH = 2.8 mmol OH^-$
 $40.0 mL \times 0.14M HOCl = 5.6 mmol HOCl$

$5.0 mL$ is $\frac{1}{2}$ Eq pt of titration therefore
 $pH = pK_a = -\log 3.2 \times 10^{-6} = \boxed{5.5}$

j $0.56M NaOH \times 7.0 mL = 3.92 mmol OH^-$



I	5.6 mmol	3.92 mmol	→	0
C	-3.92	-3.92	-	+3.92
F	1.7 mmol	0	-	+3.92 mmol
	(40.0 mL + 7 mL)			(40.0 mL + 7.0 mL)

$pH = pK_a + \log \frac{[OCl^-]}{[HOCl]}$

$5.5 + \log \left(\frac{0.083M}{0.036M} \right) = \boxed{5.9}$

$= 0.036M HOCl \quad 0.083M OCl^-$

k $15.0 mL \times 0.56M NaOH = 8.4 mmol OH^-$



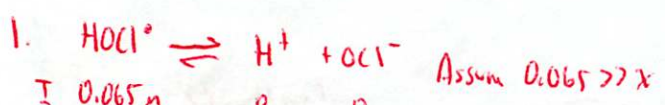
	5.6 mmol	8.4 mmol	→	0
	-5.6	-5.6	-	+5.6 mmol
		2.8 mmol	-	5.6 mmol
	55.0 mL			

$OH^- > OCl^-$ so pH is in basic is generated by OH^-

$pH = 14 - pOH$

$14 - (-\log 0.051M) =$

$= \boxed{12.74}$



I	0.065M	0	0
C	-x	+x	+x
E	0.065M - x	x	x

$\frac{x^2}{0.065M} = 3.2 \times 10^{-6} \quad [H^+] = x = 4.6 \times 10^{-4}$

$pH = \boxed{3.34}$

$\boxed{0.083M OCl^-}$ See part j



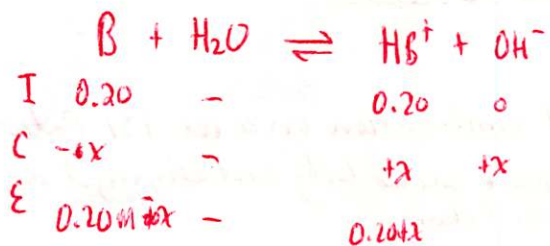
$$b \quad \% \text{ ion} = \frac{[\text{ion}]}{[\text{Initial}]} \times 100\% \quad 4.7\% = \frac{[\text{ion}]}{(0.16 \text{ M})} \times 100\%$$

$$[\text{OH}^-] = [\text{CH}_3\text{NH}_3^+] = [\text{ion}] = \boxed{7.52 \times 10^{-3} \text{ M}} \quad [\text{CH}_3\text{NH}_2] = 0.16 \text{ M} - 7.52 \times 10^{-3} \text{ M} = \boxed{0.153 \text{ M}}$$

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1 \times 10^{-14}}{7.52 \times 10^{-3}} = \boxed{1.33 \times 10^{-12} \text{ M}} \quad \text{pH} = -\log 1.33 \times 10^{-12} = \boxed{11.85}$$

$$c \quad K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} = \frac{(7.52 \times 10^{-3})^2}{0.153 \text{ M}} = \boxed{3.7 \times 10^{-4}}$$

$$d \quad [\text{CH}_3\text{NH}_2] = \frac{0.20 \text{ mol}}{1.0 \text{ L}} = 0.20 \text{ M} \quad [\text{CH}_3\text{NH}_3^+] = \frac{0.20 \text{ mol}}{1.0 \text{ L}} = 0.20 \text{ M}$$



Assume $0.20 \gg x$

$$\text{So } 0.20 \text{ M} \pm x = 0.20 \text{ M}$$

$$\frac{[\text{OH}^-](0.20 \text{ M})}{(0.20 \text{ M})} = 3.7 \times 10^{-4}$$

$$[\text{OH}^-] = 3.7 \times 10^{-4}$$

$$K_{sp} = [\text{La}^{3+}][\text{OH}^-]^3 = 1 \times 10^{-19}$$

$$\frac{0.050 \text{ mol/L La(NO}_3)_3}{1.0 \text{ L}} \times \frac{1 \text{ mol La}^{3+}}{1 \text{ mol La(NO}_3)_3} = 0.050 \text{ M La}^{3+}$$

$$Q = (0.050)(3.7 \times 10^{-4})^3 = 2.6 \times 10^{-12}$$

$$Q > K_{sp}$$

$2.6 \times 10^{-12} > 1 \times 10^{-19}$ so yes a ppt will form

3
a

$$1.2596 \text{ g Ox} \times \frac{1 \text{ mol Ox}}{126.066 \text{ g Ox}} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ox}} \times \frac{1}{41.24 \text{ mL}} \times \frac{1000 \text{ mL}}{1.0 \text{ L}} = \boxed{0.4846 \text{ M NaOH}}$$



b

$$0.4846 \text{ M NaOH} \times 38.21 \text{ mL} \times \frac{1 \text{ mol HX}}{1 \text{ mol NaOH}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.01852 \text{ mol HX}$$

$$\text{c } \frac{15.126 \text{ g}}{(5)(0.01852 \text{ mol})} = 163.3 \text{ g/mol}$$

↑
B/c 5 aliquots
of sample

d Since sample of oxalic acid would have less acid present V_{NaOH} used would be lower making M_{NaOH} higher. Higher Molarity NaOH means more moles of HX calculated, means divide in c by a bigger # so the molar mass will be decreased.

4

a Buffer solution must contain conjugate base/acid combination in a 1:1 Ratio. This solution resists change in pH when an acid or base are added by neutralizing it but since the ratio of 1:1 $\text{Acid}:\text{Base}$ will stay constant pH doesn't change.

$$\text{b } \text{pH} \approx 9 \text{ so } \text{pOH} \approx 5 \quad \text{p}K_b \approx 5 \quad K_b \approx 1 \times 10^{-5}$$

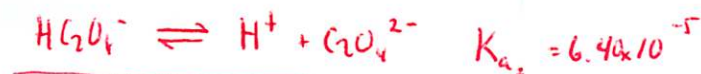
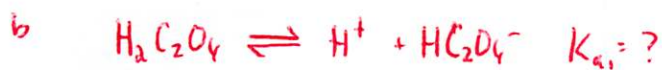
~~Debra~~ Carla C will make the correct Buffer b/c she picked a base (NH_3) w/ a K_b close to the desired K_b and she is combining the correct conj. base/acid pair ($\text{NH}_3/\text{NH}_4^+$)

Archie is wrong b/c he is mixing ~~acid~~ a combination good for pH around 5 not 9.
Beula is also wrong b/c she is mixing a combo of that will not give a buffer no conj. AB pair.
Dexter is wrong b/c he has the wrong Buffer pair when NH_3 will behave like an acid.



3

5 a $5.00 \times 10^{-3} \text{ mol } H_2A \times \frac{2 \text{ mol } OH^-}{1 \text{ mol } H_2A} \times \frac{1.00 \text{ L}}{0.400 \text{ mol } OH^-} \times \frac{1000 \text{ mL}}{1 \text{ L}} = \boxed{25.0 \text{ mL}}$



$$K_{a1} = \frac{[H^+][HA^-]}{[H_2A]}$$

$$K_w = \frac{[H]^2[A^{2-}]}{[H_2A]}$$

$$K_{a2} = \frac{[H^+][A^{2-}]}{[HA^-]}$$

$$[HA^-] = \frac{[H^+][A^{2-}]}{K_{a2}}$$

$$K_{a1} = \frac{[H^+][H^+][A^{2-}]}{[H_2A]K_{a2}} \Rightarrow K_{a1}K_{a2} = K$$

$$K_{a1} = \frac{K}{K_{a2}} = \frac{3.78 \times 10^{-6}}{6.40 \times 10^{-5}} = \boxed{0.059}$$

c pH 0.5 $[H^+] = 0.32$



I 0.15 M 0.32 0

C -x +x +x

E 0.15 M - x 0.32 + x x

Assume 0.15 M - x >> 0.15 M

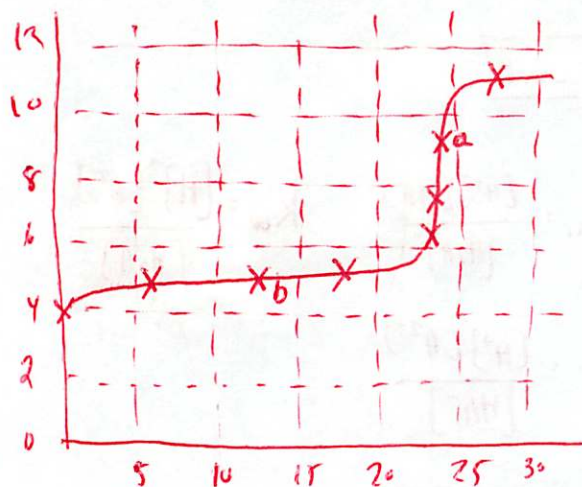
$$K = \frac{[H^+]^2[A^{2-}]}{[H_2A]}$$

$$[A^{2-}] = \frac{(0.32 \times 10^{-4})(0.15 \text{ M})}{(0.32 \text{ M})^2} = \boxed{5.5 \times 10^{-6} \text{ M}}$$



$$K_b = \frac{[HC_2O_4^-][OH^-]}{[C_2O_4^{2-}]}$$

$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{6.4 \times 10^{-5}} = \boxed{1.6 \times 10^{-10}}$$

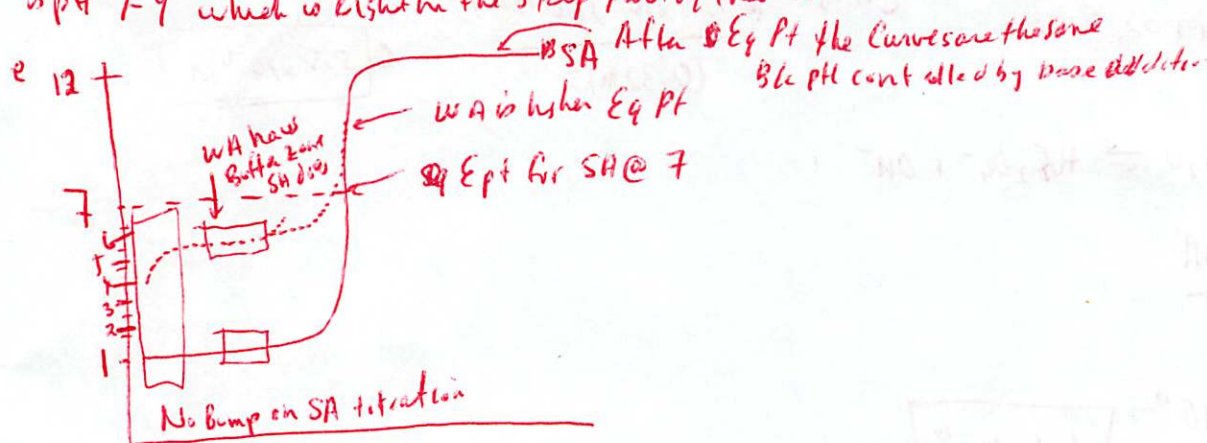


a By finding the volume need to reach the Eq Pt (x_a), at this pt the mole of H^+ added equal moles H^+ in acid sample therefore $\frac{\text{Volume NaOH} \times M_{NaOH}}{\text{Volume HCl}} = M_{HCl}$

b @ x_b is the $\frac{1}{2}$ Eq Pt where $pH = pK_a$ so take the antilog - pH @ this point to get pK_a

c $A^- + H_2O \rightleftharpoons HA + OH^-$, from this the K_b of the A^- could be found by taking the K_a of the HA acid and dividing it into the K_w (dissociation constant of H_2O)

d Cresol Red would be the best choice b/c its indicator color change is b/w $1 \times 10^{-7} - 1 \times 10^{-9}$ which is pH 7-9 which is right in the steep part of the titration curve.



$$f \quad K_b OAc^- = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.55 \times 10^{-10}$$

$$K_a HONH_3^+ = \frac{1 \times 10^{-14}}{9.1 \times 10^{-9}} = 1.1 \times 10^{-6}$$

← b/c K_a for $HONH_3^+$ is higher than K_b of OAc^- the pH of soln will be below 7 b/c the ONH_2^- will more readily donate H^+ to soln.

7 a $\text{HN}_3 \rightleftharpoons \text{H}^+ + \text{N}_3^-$

I	0.50M	0	0.082M
C	-x	+x	+x
E	0.50M-x	x	0.082M+x

Assume x is negligible

$$0.80 \text{ g NaN}_3 \times \frac{1 \text{ mol NaN}_3}{65 \text{ g NaN}_3} \times \frac{1 \text{ mol N}_3^-}{1 \text{ mol NaN}_3} \times \frac{1}{(150 \text{ mL} \times 10^{-3} \text{ L})} = 0.082 \text{ M}$$

$$K_a = \frac{[\text{H}^+][\text{N}_3^-]}{[\text{HN}_3]} = \frac{[\text{H}^+](0.082)}{(0.50 \text{ M})} = 2.8 \times 10^{-5}$$

$$[\text{H}^+] = 1.7 \times 10^{-4} \text{ M}$$

$$\text{pH} = -\log 1.7 \times 10^{-4} = \boxed{3.77}$$

b $\text{N}_3^- + \text{H}^+ \rightarrow \text{HN}_3$

I	12.3mmol 0.052	2.4mmol	75mmol
C	-2.4	-2.4	+2.4
F	9.9mmol 174mL	0	77.4mmol 174mL

24.0 mL x 0.100 M HCl = 2.4 mmol H⁺
 150 mL x 0.082 M N₃⁻ = 12.3 mmol N₃⁻
 150 mL x 0.5 M HN₃ = 75 mmol HN₃

$$\text{HN}_3 \rightleftharpoons \text{H}^+ + \text{N}_3^-$$

I	0.44M	0	0.057M
C	-x	+x	+x
E	0.44M-x	x	0.057M+x

Assume x is small

$$[\text{H}^+] = \frac{(2.8 \times 10^{-5})(0.44 \text{ M})}{0.057 \text{ M}} = 2.2 \times 10^{-4}$$

$$\text{pH} = -\log 2.2 \times 10^{-4} = \boxed{3.66}$$

c 10.0 mL x 0.1 M NaOH = 1 mmol NaOH

$$\text{HN}_3 + \text{OH}^- \rightarrow \text{N}_3^-$$

I	77.4mmol	1mmol	9.9mmol
C	-1	-1	+1
F	76.4mmol 184mL	0	10.9mmol 184mL

$$\text{HN}_3 \rightleftharpoons \text{H}^+ + \text{N}_3^-$$

I	0.43M	0	0.059M
C	-x	+x	+x
E	0.43M-x	x	0.059M+x

Assume x is small

$$[\text{H}^+] = \frac{(2.8 \times 10^{-5} \text{ M})(0.43 \text{ M})}{(0.059 \text{ M})} = 2.0 \times 10^{-4}$$

$$\text{pH} = -\log(2.0 \times 10^{-4}) = \boxed{3.69}$$